Chapter

2

Chemical Reactions and Equations

In lower class you studied about temporary, permanent , natural, man made changes etc. They may be categorized into two types known as physical changes and chemical changes. In this chapter we discuss about the chemical changes and express them in the form of chemical equations.

Consider the following processes and think about the reactions taking place during the occurance of these processes.

- coal is burnt.
- food gets digested in our body.
- iron nail is exposed to humid atmosphere for a longtime.
- we respire.
- milk is converted into curd.
- water is added to quicklime.
- crackers are burnt.

- What changes do you notice?
- Are they physical changes or chemical changes?
- Are they temporary changes or permanent changes?

In all the above processes, the nature of original substance would be changed. If new substances are formed with properties completely unlike those of the original substances, we say a chemical change has taken place.

- How do we know a chemical reaction has taken place?

Let us perform some activities to know this.

Activity 1

Take about 1 g of quick lime (calcium oxide) in a beaker. Add 10 ml of water to this. Touch the beaker with your finger.

- What do you notice?

Do you notice that the beaker is hot when you touch it. The reason is that the calcium oxide (quick lime) reacts with water and in that process heat energy is released. Calcium oxide dissolves in water producing colourless solution. Test the nature of solution with litmus paper.

- What is the nature of the solution?

A red litmus paper turns blue when dipped in the above solution. This solution is a basic solution.

Activity 2

Take about 100 ml of water in a beaker and dissolve a small quantity of sodium sulphate (Na_2SO_4).

Take about 100ml of water in another beaker and dissolve a small quantity of barium chloride ($BaCl_2$), observe the colours of the solutions obtained.

- What are the colours of the above solutions?
- Can you name the solutions obtained?

Add Na_2SO_4 solution to $BaCl_2$ solution and observe.

- Do you observe any change on mixing these solutions?

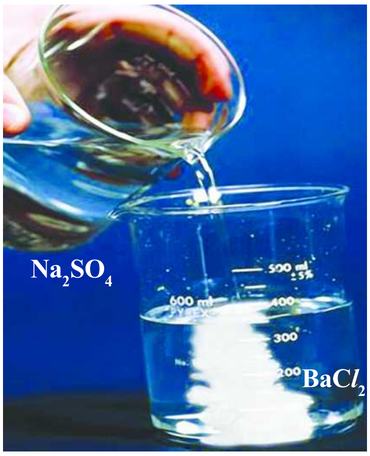

fig-1: Formation of barium sulphate precipitate

Activity 3

Take a few zinc granules in a conical flask. Add about 5 ml of dilute hydrochloric acid to the conical flask.

Observe the changes in the conical flask.

- What changes do you notice?

Now keep a burning match stick near the mouth of the conical flask.

- What happens to burning match stick?
- Touch the bottom of the conical flask with your fingers. What do you notice?
- Is there any change in temperature?

From the above activities you can conclude that during a chemical change:

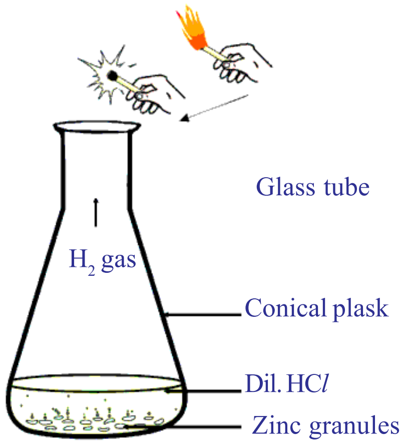

fig-2: Formation of hydrogen gas by action of dilute HCl on zinc and testing of H_2 gas

1. The original substances lose their characteristic properties. Hence these may be products with different physical states and colours.

2. Chemical changes may be exothermic or endothermic i.e, they may involve heat energy liberation or heat energy absorption.

3. They may form an insoluble substance known as precipitate.

4. There may be gas liberation in a chemical change.

In our daily life we observe variety of changes taking place around us. In this chapter we study various types of chemical reactions and their symbolic representation.

Chemical Equations

In activity 1, when calcium oxide reacts with water a new substance is formed which is unlike either calcium oxide or water. The description of chemical reactions in a sentence in activity-1 is quite long. It can be written in shorter form as a word equation.

The word equation of the above reaction is,

calcium oxide + water $\rightarrow$ calcium hydroxide .... (1)
(Reactants) (Product)

The substances which undergo chemical change in the reaction are called *reactants* and the new substances formed are called *products.*

A chemical reaction written in the form of word equation shows the change of reactants to products by an arrow placed between them. The **reactants,** are written on the left side of arrow and the final substances, or **products** are written on the right side of the arrow. The arrow head point towards the product shows the direction of the reaction.

If there is more than one reactant or product involved in the reaction, they are indicated with a plus (+) sign between them.

Writing a Chemical Equation

- Can you write a chemical reaction in any other shorter way other than the way we discussed above?

Chemical equations can be made more precise and useful if we use chemical formulae instead of words.

Generally, a compound is written by giving its **chemical formula**, which lists the symbols of the constituent elements and uses the subscript to indicate the number of atoms of each element present in the compound. If no subscript is written the number 1 is understood. Thus we can write calcium oxide as CaO, water as H_2O and the compound formed by the reaction of these two compounds is calcium hydroxide $Ca(OH)_2$.

Now the reaction of calcium oxide with water can be written as:

$CaO + H_2O \rightarrow Ca(OH)_2$ .... (2)

In the above chemical equation, count the number of atoms of each element on left side and right side of arrow.

- Is the number of atoms of each element on both sides equal?

Observe the following reactions and their chemical equations. Zinc metal reacts with dilute HCl to yield $ZnCl_2$ and liberates Hydrogen gas.

$Zn + HCl \rightarrow ZnCl_2 + H_2$ .... (3)

Sodium sulphate reacts with barium chloride to give white precipitate, barium sulphate.

$Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + NaCl$ .... (4)

- Do the atoms of each element on left side equal to the atoms of the element on the right side of the equation?

Balancing Chemical Equations

According to the law of conservation of mass, the total mass of the products formed in chemical reaction must be equal to the mass of reactants consumed. You know an atom is the smallest particle of an element that takes part in a chemical reaction. It is the atom which accounts for the mass of any substance. The number of atoms of each element before and after reaction must be the same.

All the chemical equations must balance, because atoms are neither created nor destroyed in chemical reactions. A chemical equation in which the numbers of atoms of different elements on the reactants side (left side) are same as those on product side (right side) is called a *balanced reaction*.

Balancing a chemical equation involves finding out how many **formula units** of each substance take part in the reaction. A **formula unit**, as the name implies, is one unit – whether atom, ion, or molecule – corresponding to a given formula. One formula unit of $NaCl$ for example is one Na^+ ion and one Cl^- ion; one formula unit of $MgBr_2$ is one Mg^{2+} ion and two Br^- ions and one formula unit of water is one H_2O molecule.

Now let us balance the chemical equation using a systematic method.

For this take the example of hydrogen reacting with oxygen to form water.

Step 1**:** Write the equation with the correct chemical formulae for each reactant and product.

Eg: In the reaction of hydrogen with oxygen to yield water, you can write chemical equations as follow:

$H_2 + O_2 \rightarrow H_2O$ ……(5)

Element	No of atoms	
	LHS	RHS
H	2	2
O	2	1

Step 2**:** After writing the molecular formulae of the substances the equation is to be balanced. For this we should not touch the ratio of atoms in the molecules of the susbtances but we may put suitable numbers as the coefficients before the formulae.

In the above equation put '2' before the molecular formula of hydrogen and also '2' before the molecular formula of water. Observe whether atoms of both the elements hydrogen and oxyen are same (or) different on both the sides. They are in the same number on both sides. Therefore, the equation is balanced.

$2H_2 + O_2 \rightarrow 2H_2O$ …………..(6)

Step 3**:** Sometimes there is a possibility that the coefficients of all the substances getting divided with a suitable number. Since we require the lowest ratio of coefficient for reactants and products the above operation is to be done. If no common factor is there, there is no need to divide the equation. The above equation needs no division of the coefficients of the substances.

Step 4: Verify the equation for the balancing of atoms on both sides of the equations. The above equation (6) is a balanced equation.

Let us work out some more examples to see how equations are balanced.

Eg-1: Combustion of propane (C_3H_8)

Propane, C_3H_8 is a colourless, odourless gas often used as a heating and cooking fuel. Write the chemical equation for the combustion reaction of propane. The reactants are propane and oxygen and the products are carbon dioxide and water.

Write the reaction in terms of symbols and formulae of the substances involved and follow the four steps described in previous discussion.

Step 1: Write the unbalanced equation using correct chemical formulae for all substances.

$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ ....(7), (Skeleton equation)

Note: Unbalanced chemical equation containing molecular formulae of the substances is known as *skeleton equation.*

Step 2: Compare number of atoms of each element on both sides.

Element	No of atoms	
	LHS	RHS
C	3(inC_3H_8)	1 (inCO_2)
H	8 (inC_3H_8)	2(inH_2O)
O	2 (inO_2)	3 (inCO_2H_2O)

Find the coefficients to balance the equation. It is better to start with the most complex substance – in this case C_3H_8. Look at the skeleton equation, and note that there are 3 carbon atoms on the left side of the equation but only 1 on the right side. If we add a coefficient of 3 to CO_2 on the right side the carbon atoms balance.

$C_3H_8 + O_2 \rightarrow 3CO_2 + H_2O$ .... (8)

Now, look at the number of hydrogen atoms. There are 8 hydrogens on the left but only 2 on the right side. By adding a coefficient of 4 to the H_2O on the right, the hydrogen atoms get balanced.

$C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$ ....(9)

Finally, look at the number of oxygen atoms. There are 2 on the left side but 10 on the right side, by adding a coefficient of 5 to the O_2 on the left, the oxygen atoms get balanced.

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ ....(10)

Step 3: Make sure the coefficients are reduced to their smallest whole-number values. In fact, the equation (10) is already with the coefficients in smallest whole number. There is no need to reduce its coefficients, but this might not be achieved in each chemical reaction.

Let us assume that you have got chemical equation as shown below:

$2C_3H_8 + 10O_2 \rightarrow 6CO_2 + 8H_2O$ .... (11)

- Is it a balanced equation as per rules?
- How do you say?

Though the equation (11) is balanced, the coefficients are not the smallest whole numbers. It would be necessary to divide all coefficients of equation (11) by 2 to reach the final equation.

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ .... (12)

Step 4: Check the answer. Count the numbers and kinds of atoms on both sides of the equation to make sure they are the same.

Eg-2:

Iron oxide reacts with aluminum to form iron and aluminum trioxide. Write the chemical equation to show the reaction and balance it.

Step 1: Write the equation using the correct chemical symbols and formulae for all the reactants and products.

$Fe_2O_3 + Al \rightarrow Fe + Al_2O_3$ .... (13)

Step 2: Find the suitable coefficient for the reactants and products, to equate the number of atoms of each element on both sides.

i. Examine the number of atoms of each element present in equation (13) on both sides.

Elements	No. of atoms in reactants	No. of atoms in products
Fe	2 (in Fe_2O_3)	1 (in Fe)
O	3 (in Fe_2O_3)	3 (in Al_2O_3)
Al	1 (in Al)	2 (in Al_2O_3)

In the above equation (13), Number of oxygen atoms is equal on both sides. We have to balance the remaining atoms.

ii. There are 2 Fe atoms on left side(on reactant side). There is one Fe atom on right side(on products side). To equate number of Fe atoms, multiply Fe by 2 on product side.

Now the partially balanced equation is:

$Fe_2O_3 + Al \rightarrow 2Fe + Al_2O_3$ .... (14)

iii. In the above equation(14) number of aluminum atoms still unbalanced.

There is one 'Al' atom on left side and 2 'Al' atoms on the right side (in Al_2O_3), to balance 'Al' atoms on both sides multiply 'Al' by 2 on left side of arrow mark.

Now the partially balanced equation:

$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$ .... (15)

The above equation(15), the number of atoms of each element is same on both sides of arrow mark. This is a balanced chemical equation.

Step 3: The above equation(15) is balanced and coefficients are also smallest whole numbers.

Step4: Finally to check the correctness of balanced equation, count the number of atoms of each element on both sides of equation.

$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$ .... (16)

Elements	No of Atoms of Reactants	No. of Atoms in products
Fe	2 (in Fe_2O_3)	2(in Fe)
O	3 (in Fe_2O_3)	3(in Al_2O_3)
Al	2 (in 2 Al)	2(in Al_2O_3)

(**Note:** The above method of balancing is called trial and error method only. Sometimes you may have to take more care to balance the equation.)

Making Chemical Equations more informative:

Chemical equations can be made more informative by expressing following characteristics of the reactants and products.

i. Physical state

ii. Heat changes (exothermic or endothermic change)

iii. Gas evolved (if any)

iv. Precipitate formed (if any)

i. Expressing the physical state: To make the chemical equation more informative, the physical states of the substances may be mentioned along with their chemical formulae. The different states ie., gaseous, liquid, and solid states are represented by the notations (g), (l) and (s) respectively. If the substance is present as a solution in water, the word 'aqueous' is written. In the short form it is written as (aq).

The balanced equation(16) is written along with the physical states as:

$Fe_2O_{3(s)} + 2Al_{(s)} \xrightarrow{\Delta} 2Fe_{(s)} + Al_2O_{3\,(s)}$ ...... (17) ; Δ represents heating.

fig-3(a): Aluminium in solid state

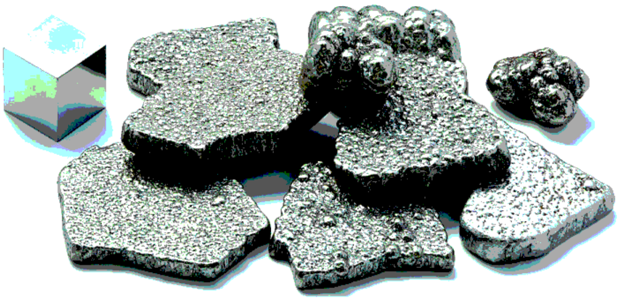

fig-3(b): Iron in solid state

ii. Expressing the heat changes: Heat is liberated in exothermic reactions and heat is absorbed in endothermic reactions. See the following examples.

1. $C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + Q$ (exothermic reaction)
2. $N_{2\,(g)} + O_{2\,(g)} \rightarrow 2NO_{(g)} - Q$ (endothermic reaction)

'Q' is *heat energy* which is shown with *plus* '+' sign on product side for exothermic reactions and *minus* '–' sign on product side for endothermic reactions.

iii. Expressing the gas evolved: If a gas is evolved in a reaction, it is denoted by an upward arrow '↑'or (g)

Eg: $Zn_{(s)} + H_2SO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + H_{2\,(g)}\uparrow$

iv. Expressing precipitate formed: If a precipitate is formed in the reactions it is denoted by a downward arrow.

Eg: $AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)}\downarrow + NaNO_{3\,(aq)}$

Sometimes the reaction conditions such as temperature, pressure, catalyst, etc are indicated above and/or below the arrow in the equation.

For example,

$$2AgCl_{(s)} \xrightarrow{\text{sunlight}} 2\,Ag_{(s)} + Cl_{2(g)}$$

$$6CO_{2\,(g)} + 6\,H_2O_{(l)} \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} \underset{\text{glucose}}{C_6H_{12}O_{6\,(s)}} + 6O_{2\,(g)}$$

Interpreting a balanced chemical equation

i. A chemical equation gives information about the reactants and products through their symbols and formulae.

ii. It gives the ratio of molecules of reactants and products.

iii. As molecular masses are expressed in 'Unified Masses' (U), the relative masses of reactants and products are known from the equation.

iv. If the masses are expressed in grams then the equation also gives the molar ratios of reactants and products.

v. If gases are involved, we can equate their masses to their volumes and calculate the volumes or those gases liberated at given condition of temperature and pressure using molar mass and molar volume relationship.

vi. Using molar mass and Avagadro's number we can calculate the number of molecules and atoms of different substances from the equation.

It gives information about relative masses of reactants and products.

from the equation we get,

a) mass - mass relationship

b) mass - volume relationship

c) volume - volume relationship

d) mass - volume - number of molecules relationship etc.,

Eg-1: $Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + Fe_{(s)}$

(atomic masses of Al = 27U, Fe = 56U, and O = 16U)

$2Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(s)}$, is a balanced equation.

(2x27)U + (2x56+3x16)U → (2x27+3x16)U + (2x56)U

54 U + 160 U → 102 U + 112 U

or 2 mol + 1 mol → 1 mol + 2 mol

54 g + 160 g → 102 g + 112 g

Suppose that your are asked to calculate the amount of aluminium, required to get 1120 kg of iron by the above reaction.

Solution: As per the balanced equation

Aluminium	→	Iron
54 g	→	112 g
x?	→	(1120 x 1000)g

$$\therefore \ x \text{ g} = \frac{(1120 \text{x} 1000)\text{g x } 54 \text{ g}}{112 \text{ g}}$$

= 10000 x 54 g

= 540000 g or 540 kg

∴ to get 1120 kg of iron we have to use 540 kg of aluminium.

Eg-2: Calculate the volume, mass and number of molecules of hydrogen liberated when 230 g of sodium reacts with excess of water at STP.(atomic masses of Na = 23U, O = 16U, and H = 1U)

The balanced equation for the above reaction is,

$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}\uparrow$

(2x23)U + 2(2x1+1x16)U → 2(23+16+1)U + (2x1)U

46 U + 36 U → 80 U + 2 U

or 46 g + 36 g → 80 g + 2 g

Solution: As per the balanced equation:

46 g of Na gives 2g of hydrogen

230g of Na gives ______? g of hydrogen.

$$\frac{230 \text{ g x } 2\text{g}}{46 \text{ g}} = 10\text{g of hydrogen}$$

1 gram molar mass of any gas at STP i.e, standard temperature 273K and standard pressure 1 bar, occupies 22.4 litres known as *gram molar volume*.

∴ 2.0g of hydrogen occupies 22.4 litres at STP.

10.0g of hydrogen occupies ....? litres at STP.

$$\frac{10.0g \times 22.4 \text{ litres}}{2.0g} = 112 \text{ litres}$$

2 g of hydrogen i.e, 1 mole of H_2 contains 6.02×10^{23} (N_O) molecules

10 g of hydrogen contain ....?

$$\frac{10.0g \times 6.02 \times 10^{23} \text{ molecules}}{2.0g}$$

$= 30.10 \times 10^{23}$ molecules

$= 3.01 \times 10^{24}$ molecules

Types of chemical reactions

In chemical reactions atoms are neither created nor destroyed. A chemical reaction is a process that is usually characterized by a *chemical change* in which the starting materials (reactants) are different from the products. Chemical reactions occur with the formation and breaking of *chemical bonds*. (you will learn about chemical bonding in chapter …) Some common reaction types are discussed below.

Chemical Combination

Activity 4

(This activity needs Teacher's assistance)

- Take a small piece (about 3 cm long) of magnesium ribbon.
- Rub the magnesium ribbon with sand paper.
- Hold it with a pair of tongs.
- Burn it with a spirit lamp or burner.

• What you observe?

You will notice that,

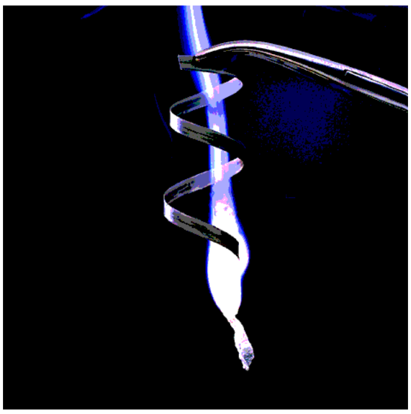

fig-4: Burning of magnesium ribbon

Magnesium burns in oxygen by producing dazzling white flame and changes into white powder. The white powder is magnesium oxide.

$$\underset{\text{Magnesium}}{2Mg_{(s)}} + \underset{\text{Oxygen}}{O_{2\,(g)}} \rightarrow \underset{\text{Magnesium oxide}}{2MgO_{(s)}} \quad \text{......... (18)}$$

In this reaction magnesium and oxygen combine to form a new substance magnesium oxide. A reaction in which single product is formed from two or more reactants is known as *chemical combination reaction*.

You will also notice release of enormous amount of heat energy when magnesium is burnt in air.

Let us discuss some more examples of combination reactions.

i. Burning of Coal: When coal is burnt in oxygen, carbon dioxide is produced.

$C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + Q$ (heat energy) ....(19)

ii. Slaked lime is prepared by adding water to quick lime.

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2\,(aq)} + Q$ (heat energy).....(20)

Large amount of heat energy is released on reaction of water with $CaO_{(s)}$. If you touch the walls of the container you will feel the hotness. Such reactions are called **exothermic** reactions.

fig-5: Formation of slaked lime by the reaction of CaO with water

A solution of slaked lime produced in the reaction equation(20) is used to white wash the walls. Calcium hydroxide reacts slowly with the carbon dioxide in air to form a thin layer of calcium carbonate on the walls. It gives a shiny finish to the walls.

$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$

The chemical formula of marble is $CaCO_3$

Decompostion Reaction

Activity 5

- Take a pinch of calcium carbonate (lime stone) in a boiling tube.
- Heat the boiling tube over the flame of spirit lamp or burner.
- Now bring a burning match stick near the evolved gas as shown in the figure.
- What do you observe?

You will notice that match stick would be put off.

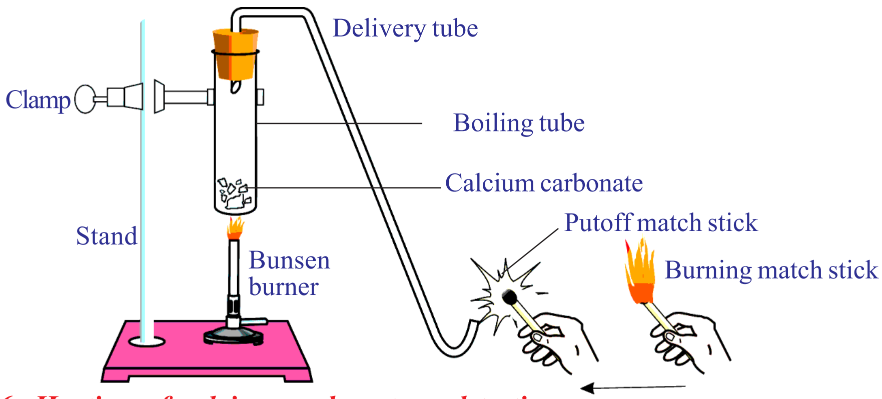

fig-6: Heating of calcium carbonate and testing the gas evolved with burning match stick

In the above activity, on heating calcium carbonate decomposes to calcium oxide and carbon dioxide.

$$\underset{\text{Lime stone}}{CaCO_{3\,(s)}} \xrightarrow{\text{Heat}} \underset{\text{quick lime}}{CaO_{(s)}} + CO_{2\,(g)} \;\ldots\ldots\ldots\ldots\ldots\ldots (21)$$

It is a thermal decomposition reaction. When a decomposition reaction is carried out by heating, it is called *thermal decomposition reaction*.

Activity 6

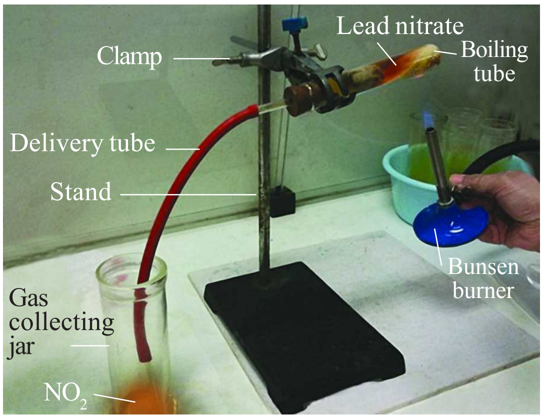

fig-7:Heating of lead nitrate and emission of nitrogen dioxide

- Take about 0.5g of lead nitrate powder in a boiling test tube.
- Hold the boiling tube with a test tube holder.
- Heat the boiling tube over a flame. (see figure)
- Note down the change.
- What do you observe?

Heating of lead nitrate and emission of nitrogen dioxide

On heating lead nitrate decomposes to lead oxide, oxygen and nitrogen dioxide. You observe the brown fumes liberating in the boiling tube. These brown fumes are of nitrogen dioxide (NO_2).

$$\underset{\text{Lead Nitrate}}{2Pb\,(NO_3)_{2(s)}} \xrightarrow{\text{Heat}} \underset{\text{lead oxide}}{2PbO_{(s)}} + \underset{\text{Nitrogen dioxide}}{4NO_{2\,(g)}} + \underset{\text{Oxygen}}{O_{2\,(g)}} \;\ldots\ldots\ldots (22)$$

This is also a thermal decomposition reaction.

Let us perform some more decomposition reactions

Activity 7

- Take a plastic mug. Drill two holes at its base.
- Fit two 'one holed rubber stoppers' in these holes.
- Insert two carbon electrodes in these rubber stoppers.
- Connect the electrodes to 9V battery as shown in fig.
- Fill the mug with water, so that the electrodes are immersed.
- Add few drops of dilute sulphuric acid to water.
- Take two test tubes filled with water and invert them over the two carbon electrodes.
- Switch on the current and leave the apparatus undisturbed for some time.

• What do you observe in the test tubes?

You will notice the liberation of gas bubbles at both the electrodes. These bubbles displace the water in the test tubes.

Is the volume of gas collected in both the test tubes same?

Once the test tubes are filled with gases take them out carefully. Test both the gases separately by bringing a burning candle near the mouth of each test tube.

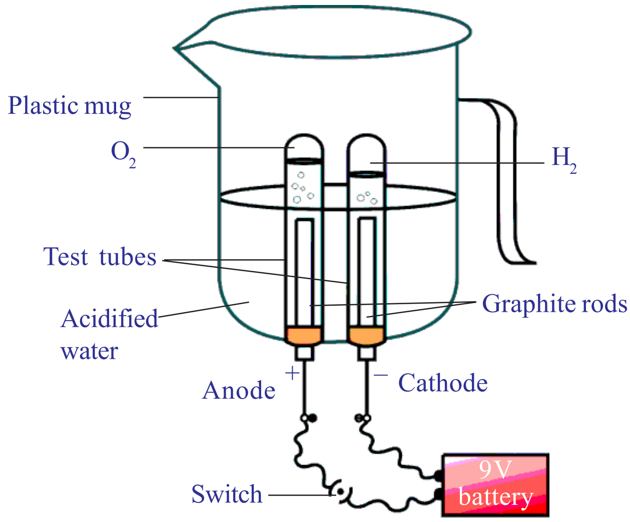

fig-8: Electrolysis of water

- What do you observe in each case?

Can you predict the gas present in each test tube?

In the above activity on passing the electricity, water dissociates to hydrogen and oxygen.

$$2H_2O_{(l)} \xrightarrow{\text{Electrolysis}} 2\,H_{2\,(g)} + O_{2\,(g)} \quad .... (23)$$

Activity 8

- Take some quantity of silver bromide on a watch glass.
- Observe the colour of silver bromide.
- Place the watch glass in sunlight for some time.
- Now observe the colour of silver bromide.

- What changes do you notice?
- Did the colour of the silver bromide change?

fig-9(a): Silver bromide (light yellow colour)

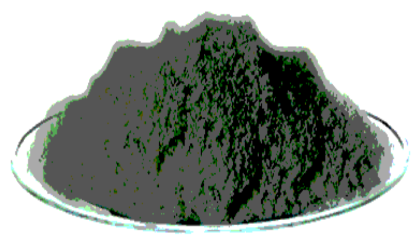

fig-9(b) when exposed to sunlight (gray colour) silver metal

Silver bromide decomposes to silver and bromine in sunlight. *Light yellow* coloured silver bromide turns to *gray* due to sunlight.

$$2AgBr_{(s)} \xrightarrow{\text{sunlight}} 2\,Ag_{(s)} + Br_{2\,(g)} \qquad \text{.............................. (24)}$$

This decomposition reaction occurs in presence of sunlight and such reactions are called *photochemical reactions*.

All the above decomposition reactions require energy in the form of heat, light or electricity for converting the reactants to products. All these reactions are endothermic.

Carry out the following Activities:

i) Take a pinch of $AgCl_2$ in a watch glass. Keep it in sunlight for some time and observe the change.

ii) Take some ferrous sulphate crystals in a boiling tube. Heat it over spirit lamp.

iii) Take about 2 gm of barium hydroxide in a test tube. Add about 1 gm of ammonium chloride and mix with glass rod. Touch the test tube with your palm.

What do you observe?

Displacement reaction.

In displacement reaction one element displaces another element from its compound and takes its place there in.

Displacement of hydrogen from acids by metals:

Generally metals which are more active than hydrogen displace it from an acid.

Let us observe the reaction in following activity.

Activity 9

- Take a small quantity of zinc dust in a conical flask.
- Add dilute hydrochloric acid slowly.
- Now take a balloon and tie it to the mouth of the conical flask.
- Closely observe the changes in the conical flask and balloon.
- What do you notice?

You can see the gas bubbles coming out from the solution and the balloon bulges out (figure 10b). Zinc pieces react with dilute hydrochloric acid and liberate hydrogen gas as shown below.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2\,(g)} \qquad \text{..........................(25)}$$

In reaction (25) the element zinc has displaced hydrogen from hydrochloric acid. This is displacement reaction.

fig-10(a)

fig-10(b)

Activity 10

- Take two iron nails and clean them by rubbing with sand paper.
- Take two test tubes and mark them as A and B.
- Take about 10ml of copper sulphate solution in each test tube. Dip one iron nail in copper sulphate solution of test tube A and keep it undisturbed for 20 minutes.
- Keep the other iron nail aside.
- Now take out the iron nail from copper sulphate solution and compare with the other iron nail that has been kept aside. (see fig11-a)
- Compare the colours of the solutions in the test tubes. (see fig11-b)
- What changes do you observe?

You will find the iron nail dipped in copper sulphate solution becoming brown. The blue colour of copper sulphate solution in test tube 'A' fades.

fig-11(a):Iron nail dipped in copper sulphate solution

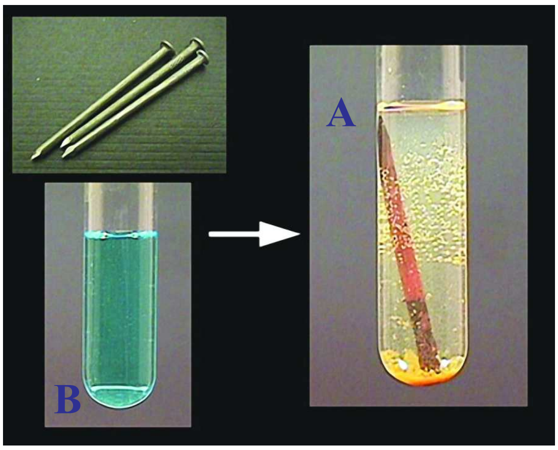

fig-11(b): Iron nail and copper sulphate solutions compared before and after the experiment

The chemical reaction in this activity is:

$$Fe_{(s)} + CuSO_{4\,(aq)} \rightarrow FeSO_{4\,(aq)} + Cu_{(s)} \quad ....(26)$$

Iron is more reactive than copper, so it displaces copper from copper sulphate. This is another example of displacement reaction.

Other examples of displacement reaction are:

i) $Zn_{(s)} + 2AgNO_{3\,(aq)} \rightarrow Zn(NO_3)_{2\,(aq)} + 2Ag_{(s)}$ .... (27)

ii) $Pb_{(s)} + CuCl_{2\,(aq)} \rightarrow PbCl_{2\,(aq)} + Cu_{(s)}$ .... (28)

Double displacement reaction

Activity 11

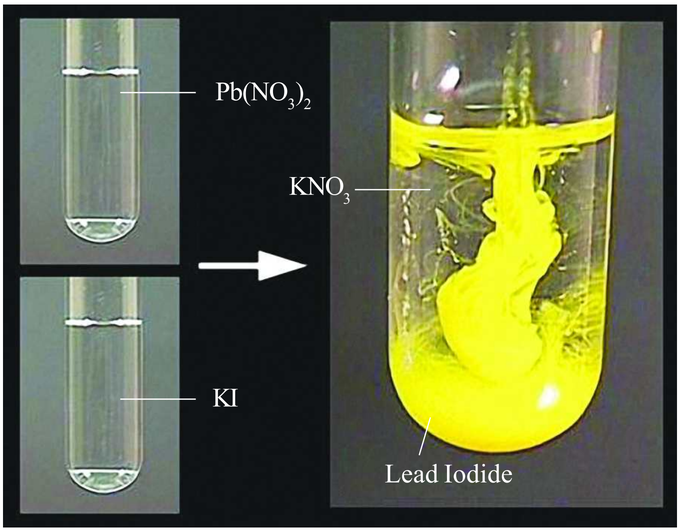

fig-12: formation of lead iodide and potassium nitrate

- Take a pinch of lead nitrate and dissolve in 5.0ml of distilled water in a test tube.
- Take a pinch of potassium iodide in a test tube and dissolve in distilled water.
- Mix lead nitrate solution with potassium iodide solution.
- What do you observe?

A yellow coloured substance which is insoluble in water, is formed as **precipitate**. The precipitate is lead iodide.

$$\underset{\text{lead nitrate}}{Pb(NO_3)_{2(aq)}} + \underset{\text{potassium Iodide}}{2KI_{(aq)}} \rightarrow \underset{\text{lead iodide}}{PbI_{2\,(s)}} + \underset{\text{potassium nitrate}}{2KNO_{3\,(aq)}} \quad ....(29)$$

This reaction is double displacement reaction. If two reactants exchange their constituents chemically and form two products, then the reaction is called as double displacement reaction.

Other examples of double displacement reactions are:

1) Sodium sulphate solution on mixing with barium chloride solution forms a white prepitate of barium sulphate and soluble sodium chloride.

$$Na_2SO_{4\,(aq)} + BaCl_{2\,(aq)} \rightarrow BaSO_{4(s)} + 2\,NaCl_{(aq)} \quad ....(30)$$

2) Sodium hydroxide reacts with hydrochloric acid to form sodium chloride and water.

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$ ....(31)

3) Sodium chloride spontaneously combines with silver nitrate in solution giving silver chloride precipitate.

$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$ ....(32)

Oxidation and Reduction

'Oxidation' is a reaction that involves the addition of oxygen or removal of hydrogen.

'Reduction' is a reaction that involves the addition of hydrogen or removal of oxygen.

Let us try to understand more clearly doing this experiment.

Activity 12

- Take about 1.0g of copper powder in a china dish.
- Keep the china dish on a tripod stand containing wire gauge.
- Heat it with a bunsen burner or with a spirit lamp.
- Do you find any change in colour of copper?

You will notice that the surface layer of copper becomes black.

- Why does the colour of copper change?
- What is the black colour product formed on the surface of copper?

In the activity on heating copper it reacts with oxygen present in the atmosphere to form copper oxide.

The reaction is shown below.

$Cu_{(s)} + O_{2(g)} \xrightarrow{Heat} 2\,CuO_{(S)}$ ....(33)

fig-13(a): Black copper oxide

fig-13(b) China dish

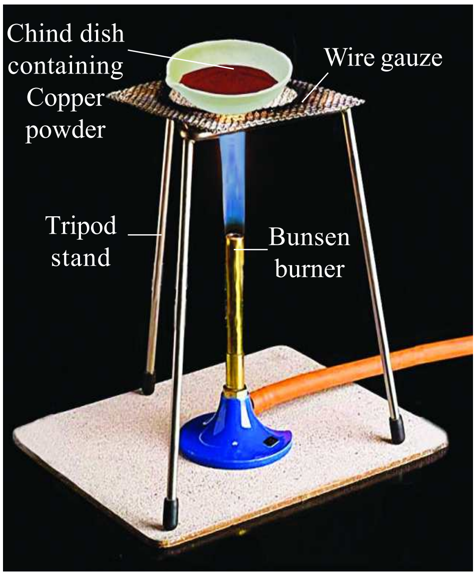

fig-13(c): Oxidation of copper to copper oxide

Here copper combines with oxygen to form copper oxide. Here oxygen is gained and the process is called *oxidation*.

Now pass hydrogen gas over hot copper oxide obtained in above activity and observe the change.

- What do you notice?
- Is there any change in black colour of copper oxide?

You will notice that the black coating on copper turns brown because copper oxide loses oxygen to form copper. In this process oxygen is lost and the process is called *Reduction*.

$$CuO_{(s)} + H_{2(g)} \xrightarrow{\text{Heat}} Cu_{(s)} + H_2O_{(g)}$$

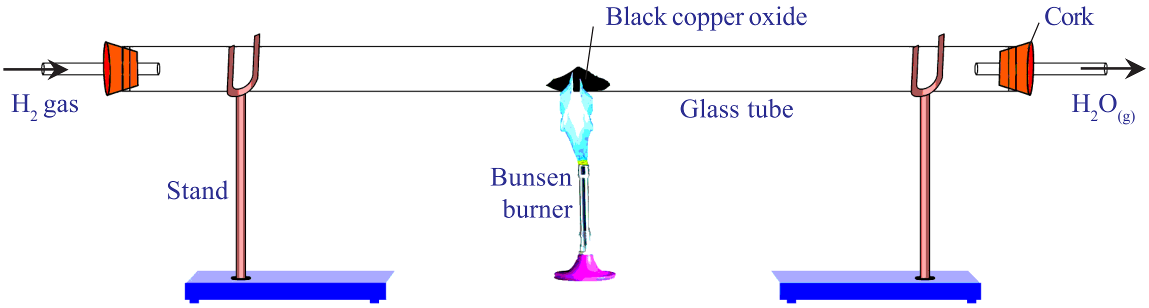

fig-14: Reduction of copper oxide to copper

In the above reaction hydrogen is gained; such reaction is called *reduction reaction*.

Generally oxidation and reduction occur in the same reaction. If one reactant gets oxidized, the other gets reduced. Such reactions are called *oxidation-reduction reactions or redox reactions*.

In the CuO, H_2 reaction CuO is reduced and H_2 is oxidized.

Some other examples of redox reactions are:

i) $2\,Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4\,Fe_{(s)} + 3CO_{2\,(g)}$ .... (34)

ii) $2\,PbO_{(3)} + C_{(s)} \rightarrow 2Pb_{(s)} + CO_{2\,(g)}$ .... (35)

Have you observe the effects of oxidation reactions in daily life

Corrosion:

You must have observed that a freshly cut apple turns brown after some time. The shining iron articles gradually become reddish brown when left for some time. Burning of crackers produce dazzling light with white fumes.

- How do these changes occur?

They are all the examples of the process called *oxidation.*

Let us know how?

Oxidation is the reaction of oxygen molecules with different substances starting from metal to living tissue which may come in contact with it.

fig-15: Rusting of iron

Apples pears, bananas, potatoes etc., contain enzyme called polyphenol oxidase or tyrosinase, which reacts with oxygen and changes the colour on the cut surface of the fruit.

The browning of iron, when left for sometime in moist air, is a process commonly known as rusting of iron. This process is basically oxidation reaction which requires both oxygen and water. Rusting does not occur in oxygen free water or dry air.

Burning of crackers is also oxidation process of variety of chemicals.

- Did you notice the colour coating on silver and copper articles?

When some metals are exposed to moisture, acids etc., they tarnish due to the formation of respective metal oxide on their surface. This process is called ***corrosion***.

Look at the following examples:

i. The black coatings on silver(see fig-16)

fig-16: Tarnishing of silverware (before & after)

$$4Ag + 2H_2S + O_2 \rightarrow \underset{\text{black}}{2Ag_2S} + 2H_2O \quad \text{..............} (37)$$

ii. Green coating on copper (see fig-17)

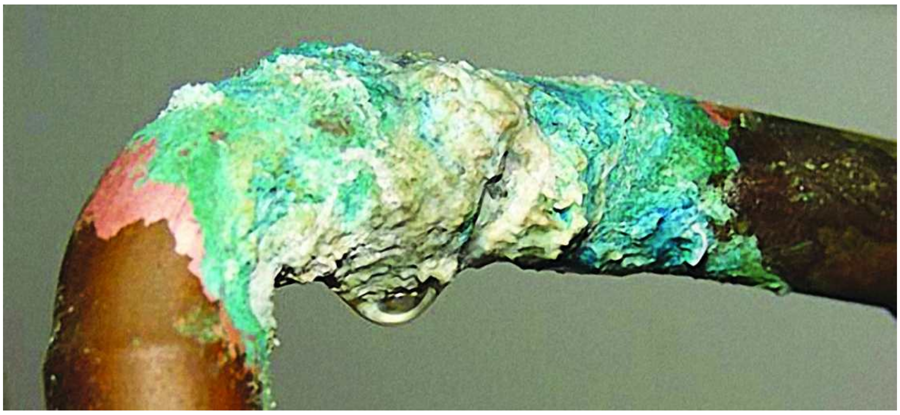

fig-17: Corrosion of copper

$2Cu + O_2 \rightarrow 2CuO$ .... (38)

Corrosion is the oxidative deterioration of a metal.

Corrosion causes damage to car bodies, bridges, iron railings, ships etc., and to all other objects that are made of metals. Especially corrosion of iron is a serious problem.

Corrosion can be prevented or at least minimized by shielding the metal surface from oxygen and moisture. It can be prevented by painting, oiling, greasing, galvanizing, chrome plating or making alloys. *Galvanizing* is a method of protecting iron from rusting by coating them a thin layer of Zinc.

Alloying is also a very good method of improving properties of metal. Generally pure form of iron is very soft and stretches easily when hot. Iron is mixed with carbon, nickel and chromium to get an alloy stainless steel. The stainless steel is hard and does not rust.

A metallic substance made by mixing and fusing two or more metals, or a metal and a nonmetal, to obtain desirable qualities such as hardness, lightness, and strength is known as *alloy*.

For example: Brass, bronze, and steel.

Do you know?

Gold one of the most valuable of elements has been prized since antiquity for its beauty and resistance to corrosion.

Some more effects of oxidation on everyday life

- Combustion is the most common example for oxidation reactions.

  For example: burning of wood involves release of carbon dioxide, water vapour along with huge amount of energy.
- Rising of dough with yeast depends on oxidation of sugars to carbon dioxide and water.
- Bleaching of coloured objects using moist chlorine

$$Cl_2 + H_2O \rightarrow HOCl + HCl \qquad \text{..............} (39)$$

$$HOCl \rightarrow HCl + (O) \qquad \text{..............} (40)$$

Coloured object + (O) $\rightarrow$ Colourless object.

Some times during rainy season the power supply to our home from the electric pole will be interrupted due to formation of the metal oxide layer on the electric wire. This metal oxide is an electrical insulator. On removing the metal oxide layer formed on the wire with a sand paper, supply of electricity can be restored.

Rancidity

- Have you ever tasted or smelt the fat/oil containing food materials left for a long time?
- When fats and oils are oxidized they become rancid. Their smell and taste change.

Thus we can say that oxidation reactions in food material that were left for a long period are responsible for spoiling of food.

Rancidity is an oxidation reaction.

- How can we prevent the spoiling of food?

The spoilage of food can be prevented by adding preservatives like Vitamin C and Vitamin E.

Usually substances which prevent oxidation (Antioxidants) are added to food containing fats and oil. Keeping food in air tight containers helps to slow down oxidation process.

Do you know that manufacturers of potato chips flush bags of chips with nitrogen gas to prevent the chips from getting oxidized.

Key words

Reactants, products, exothermic reaction, endothermic reaction, chemical combination, chemical decomposition, displacement reaction, double displacement reaction, oxidation, reduction, corrosion, rancidity, antioxidants.

What we have learnt

- Chemical change is permanent change.
- A Chemical equation represents a chemical reaction.
- Complete chemical equation represents the reactants, products and their physical state.
- A Chemical equation is said to be balanced, when the number of atoms of each element is same on both reactants side and products side.
- A chemical equation must always be balanced.
- In a combination reaction two or more substances combine to form a new single substance.
- In a decomposition reaction a single substance decomposes to give two or more substances.
- Reactions in which heat energy is absorbed by the reactants are endothermic reactions.
- In exothermic reaction heat energy is released by the reactants.
- A displacement reaction occurs, when an element displaces another element from its compound.
- Two different atoms or ions are exchanged in double displacement reactions.
- Oxidation is the gain of Oxygen or loss of Hydrogen.
- Loss of oxygen or gain of Hydrogen is Reduction.
- Corrosion causes damage to iron appliances.
- When fats and oils are oxidized, they become rancid.
- Precipitate is an insoluble substance.

Improve your learning

1. What is a balanced chemical equation? Why should chemical equations be balanced? (AS1)
2. Balance the following chemical equations. (AS1)
   a) $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
   b) $Hg(NO_3)_2 + KI \rightarrow HgI_2 + KNO_3$
   c) $H_2 + O_2 \rightarrow H_2O$
   d) $KClO_3 \rightarrow KCl + O_2$
   e) $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
3. Write the balanced chemical equations for the following reactions. (AS1)
   a) Zinc + Silver nitrate → Zinc nitrate + Silver.
   b) Aluminum + copper chloride → Aluminum chloride + Copper.
   c) Hydrogen + Chlorine. → Hydrogen chloride.
   d) Ammonium nitrate → Nitrogen + Carbon dioxide + water.
4. Write the balanced chemical equation for the following and indentify the type of reaction in each case. (AS1)
   a) Calcium hydroxide $_{(aq)}$ + Nitric acid $_{(aq)}$ → Water $_{(l)}$ + Calcium nitrate $_{(aq)}$

b) Magnesium $_{(s)}$ + Iodine $_{(g)}$ → Magnesium Iodide. $_{(s)}$

c) Magnesium$_{(s)}$ + Hydrochloric acid$_{(aq)}$ → Magnesium chloride$_{(aq)}$ + Hydrogen$_{(g)}$

d) Zinc$_{(s)}$ + Calcium chloride $_{(aq)}$ → Zinc Chloride $_{(aq)}$ + Ca$_{(s)}$

5. Write an equation for decomposition reaction where energy is supplied in the form of Heat / light / electricity. (AS1)
6. What do you mean by precipitation reaction? (AS1)
7. How chemical displacement reactions differ from chemical decomposition reaction? Explain with an example for each. (AS1)
8. Name the reactions taking place in the presence of sunlight? (AS1)
9. Why does respiration considered as an exothermic reaction? Explain. (AS1)
10. What is the difference between displacement and double displacement reactions? Write equations for these reactions? (AS1)
11. $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

    In the above equation, name the compound which is oxidized and which is reduced? (AS1)
12. Give two examples for oxidation-reduction reaction. (AS1)
13. In the refining of silver, the recovery of silver from silver nitrate solution involved displacement by copper metal. Write the reaction involved. (AS1)
14. What do you mean by corrosion? How can you prevent it? (AS1)
15. Explain rancidity. (AS1)
16. Balance the following chemical equations including the physical states. (AS1)

    a) $C_6H_{12}O_6 \longrightarrow C_2H_5OH + CO_2$

    b) $Fe + O_2 \longrightarrow Fe_2O_3$

    c) $NH_3 + Cl_2 \longrightarrow N_2H_4 + NH_4Cl$

    d) $Na + H_2O \longrightarrow NaOH + H_2$
17. Balance the chemical equation by including the physical states of the substances for the following reactions. (AS1)

    a) Barium chloride and sodium sulphate aqueous solutions react to give insoluble Barium sulphate and aqueous solution of sodium chloride.

    b) Sodium hydroxide reacts with hydrochloric acid to produce sodium chloride and water.

    c) Zinc pieces react with dilute hydrochloric acid to liberate hydrogen gas and forms zinc chloride
18. A shiny brown coloured element 'X' o n heating in air becomes black in colour. Can you predict the element 'X' and the black coloured substance formed? How do you support your predictions? (AS2)
19. Why do we apply paint on iron articles? (AS7)
20. What is the use of keeping food in air tight containers? (AS7)

Fill in the blanks

1. The decomposition of vegetable into compost is an example of ____________ reaction.
2. The chemical reactions in which energy is absorbed to form a new compound is called ________
3. The reaction $2N_2O \rightarrow 2N_2 + O_2$ is an example for ________ reaction.
4. The reaction $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2\uparrow$ is an example for _______ reaction.
5. The substances that are present on left side of a chemical equation are called _____.
6. The arrow mark between the products and reactants of a chemical equation shows ______ of the reaction.
7. **Match the following:**

1) $2AgNO_3 + Na_2CrO_4 \rightarrow Ag_2CrO_4 + 2NaNO_3$	( )	a) combination reactions
2) $2NH_3 \rightarrow N_2 + 3H_2$	( )	b) decomposition reactions
3) $C_2H_4 + H_2O \rightarrow C_2H_6O$	( )	c) displacement reactions
4) $Fe_2O_3 + 3CO \rightarrow 2F_2 + 3CO_2$	( )	d) double displacement Reactions

Multiple choice questions

1. $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$. [ ]
   The above reaction is an example of:
   a) Combination reaction   b) Decomposition reaction
   c) Displacement reaction   d) Double decomposition reaction
2. What happens when dil. hydrochloric acid is added to iron filings? Choose the correct answer.
   a) Hydrogen gas and iron chloride are produced. [ ]
   b) Chlorine gas and iron hydroxide are produced.
   c) No reaction takes place.
   d) Iron salt and water are produced.
3. $2PbO_{(s)} + C_{(s)} \rightarrow 2Pb_{(s)} + CO_{2(g)}$ [ ]
   Which of the following statements are correct for the above chemical reaction?
   a) Lead is reduced   b) Carbon dioxide is oxidized
   c) Carbon is oxidized   d) Lead oxide is reduced.
   i) (a) and (b)   ii) (a) and (c)   iii) (a), (b) and (c)   iv) all.
4. The chemical equation $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$ represents following type of chemical reaction. [ ]
   i) displacement   ii) combination
   iii) decomposition   iv) double-displacement
5. The reaction of formation hydrogen chloride from hydrogen and chloride represents following type of chemical reaction [ ]
   i) decomposition   ii) displacement
   iii) combination   iv) double-displacement